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Unit 2: atoms, molecules and stoichiometry

2.1 Relative masses of atoms and molecules

Atomic mass unit: one twelfth the mass of a carbon-12 atom

Relative atomic mass: the weighted average mass of an atom relative to 1/12 of the mass of a carbon-12 atom

Relative isotopic mass: the weighted average mass of an atom of an isotope relative to 1/12 the mass of a carbon-12 atom

Relative molecular mass: the weighted average mass of a molecule relative to 1/12 the mass of a carbon-12 atom

Relative formula mass: the weighted average mass of the atoms in a formula relative to 1/12 the mass of a carbon-12 atom

$$A_r = \frac{\Sigma(isotope\ abundance\ \times isotope\ mass\ number)}{100}$$

2.2 The mole and the Avogadro constant

Mole: of a substance is the amount of that substance that contains the Avogadro constant number of specified particles

Avogadro constant: 6.02 x 10^{23} molecules

2.3 Formulae

Ions required in our syllabus that can't be inferred from periodic table:

Nitrate (NO_3^-), Carbonate (CO_3^{2-}), sulfate (SO_4^{2-}), Hydroxide (OH^-), ammonium (NH_4^+), zinc ion (Zn^{2+}), silver ion (Ag^+), Hydrogen carbonate (HCO_3^-), Phosphate (PO_4^{3-})

Common states: ions are always aqueous, soluble salts are usually aqueous, precipitates are solids, acids and alkalis are always aqueous

Empirical formula: the simplest whole number ratio of the elements present in one molecule or formula unit of the compound.

Molecular formula: the actual number of each of the different atoms of each element present in a molecule.

Anhydrous substance: doesn't contain water

Hydrated substance: the element of water is added to it

Water of crystallization: water which is locked into a crystal in a fixed way

How to calculate molecular formula: count number and type of atoms and put them in a formula

How to calculate empirical formula: the simplest form of all the co-efficient for each type of atom, or if given masses:

Finding empirical formulae from mass data

Suppose you found that 0.46 g of sodium formed 0.78 g of sodium sulfide. That means that 0.46 g of sodium combines with (0.78 - 0.46) g = 0.32 g of sulfur.

It is clearest if you set your answer out as a simple table:

	Na	S
mass	0.46 g	0.32 g
number of moles of atoms	0.46/23	0.32/32
	= 0.02	= 0.01
ratio	2	1

If the compound only contains carbon and hydrogen

A compound which only contains carbon and hydrogen is called a hydrocarbon. Look for this word in the question. You can only use this shorter method if you *know* that there is nothing else present.

Here's an example:

When 0.78 g of a hydrocarbon was burned in excess air, 2.64 g of carbon dioxide and 0.54 g of water were formed. Find the empirical formula of the hydrocarbon.

The important things to notice is that every mole of CO_2 contains 1 mole of carbon atoms. Every mole of H_2O contains 2 moles of hydrogen atoms.

1 mole of CO_2 weighs 44 g.

No of moles of CO_2 = 2.64/44 = 0.06

Therefore, no of moles of carbon atoms, C = 0.06

1 mole of H_2O weighs 18 g.

No of moles of H_2O = 0.54/18 = 0.03

Each mole of H_2O contains 2 moles of hydrogen atoms.

Therefore, no of moles of hydrogen atoms, H = 0.06

The ratio of the number of moles of C : H is 1 : 1.

The empirical formula is CH.

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2.4 Reacting masses and volumes

Moles = mass/mR

Moles = volume x concentration

Moles = volume/24dm^3

Conversions:

1 m = 10dm

1m = 100 cm

1m^2 = 100dm^2

1dm^3 = 1000cm^3

Types of reactions:

Displacement reactions – where a more reactive element takes the place of a less reactive element in a compound

Acid base reactions – where an acid and a base react to form salt and water

Precipitation reaction – where solid is produced when 2 aqueous reactants react

More about this topic in paper 3 notes

Extra page for pastpapers notes